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ABSTRACT

Introduction to Engineering Chemistry is a four-credit hour (one semester) course designed to introduce marine engineering students to the rudiments of basic (introductory) inorganic chemistry. The course consists of 18 units (numbered 1.0 through 18.0) focusing on these subject areas: fundamental concepts; structure of the atom and the periodic law; the relationship of periodic classification to the properties of the elements; chemical bonding; chemical symbols, formulas, and equations; the gaseous state and kinetic molecular theory; liquid and solid states; solutions; oxidation-reduction reactions; acids, bases, and salts; chemical equilibria; ionic equilibria involving weak electrolytes; solubility product principle; electrochemistry; chemistry and properties of liquified natural gas; alloys; nuclear chemistry; and corrosion. Terminal objectives (which take on the number of the units followed by successive units in the tenths place) and enabling objectives (which take the number of the unit and the number of the terminal objective followed by successive units in the hundredths place) are provided for each of the 18 units. (JN)

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CURRICULUM OUTLINE

for

INTRODUCTION TO ENGINEERING CHEMISTRY

(Second Edition)

(Review Cycle-Annual)

by

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INTRODUCTION: Curriculum outlines provide instructors, among other things, with behaviors students should be able to demonstrate when they have completed a course, a unit of instruction, or a lesson. These behaviors, called terminal and enabling objectives, are the objectives over which students, in one way or another, are tested. Any testing should determine whether students can demonstrate minimum levels of competency with the objectives set forth by the instructors. Terminal objectives are somewhat broad in scope when compared with enabling objectives. Additionally, enabling objectives are subsets of terminal objectives; each terminal objective has several enabling objectives. Enabling objectives are the objectives which allow teachers to decide to what degree their students are competent with a terminal objective. To be competent with a terminal objective, however, does not mean a student must demonstrate competency with every enabling objective of a specific terminal objective. The actual number of enabling objectives over which students must show facility is set by the course instructor and constitutes a performance standard or criterion. It may be, for example, that the minimum competency level for a certain terminal objective is 70%. This may translate to students having to show facility with 70% of the enabling objectives for that terminal objective. The competency level may change from one to another terminal objective.

The curriculum outline has been divided into units, each numbered successfully from 1.0 to the end. Each unit may involve one or more lectures. The number of lessons is a matter of instructor style, desires, and the like. Terminal objectives take on the numbers of the units, 1 to as many as there are, followed by successive numerals in the tenths place (for example, 1.1, 1.2, 1.3). In the example, unit 1 has 3 terminal objectives. Enabling objectives take on

the number of the unit, then the number of the terminal objective followed by the numbering in the hundredths place from 1 to the end (for example, 1.1.1, 1.1.2, 1.1.3, 2.1.1, 2.1.2, 2.2.1, 2.2.2).

MISSION: Introduction to Engineering Chemistry is a four semester hour (one semester) post secondary course. The course includes four one-hour lectures each week. Assumptions are not made regarding the background of entering students. High school chemistry is not a prerequisite for this course. The mission of the course is to introduce marine engineering students to the rudiments of basic or introductory inorganic chemistry. Special topics, when included, are engineering oriented to the extent the theory involved may, with a relative degree of ease, be applied to the marine world of shipping and further the marine environment and its effects upon ship hulls.

SCOPE: The following subjects are discussed in Introduction to Engineering Chemistry; (1) fundamental concepts; (2) structure of the atom and the periodic law; (3) the relationship of periodic classification to the properties of the elements; (4) general principles of chemical bonding; (5) chemical symbols, formulae and equations; (6) the gaseous state and kinetic molecular theory; (7) liquid and solid states; (8) solutions; (9) oxidation-reduction reactions; (10) acids, bases and salts; (11) chemical equilibra; (12) ionic equilibra involving weak electrolytes; (13) solubility product principle; (14) electrochemistry; (15) chemistry and properties of liquified natural gas; (16) alloys; (17) nuclear chemistry; (18) corrosion.

TERMINAL AND ENABLING OBJECTIVES: Upon completion of this course, students should:

UNIT 1.0 Fundamental Concepts

1.1 Given questions and problems DEMONSTRATE knowledge of the fundamental concepts of science relevant to introductory chemistry to a 60% level of competence.

1.1.1 DEFINE matter and energy.

1.1.2 DESCRIBE the relationships between matter and energy.

1.1.3 DIFFERENTIATE between three states of matter.

1.1.4 DISTINGUISH between elements, compounds and mixtures.

1.1.5 DISTINGUISH between the physical and chemical properties of a substance.

1.1.6 USE the metric units when solving problems.

1.1.7 LIST the basic metric units and their common prefixes.

1.1.8 CONVERT temperatures between the Fahrenheit, celsius and Kelvin scales.

1.1.9 PERFORM calculations involving density.

1.1.10 DEFINE specific gravity and specific heat.

1.1.11 IDENTIFY the differences between chemical and physical changes in matter.

UNIT 2.0 Structure of the Atom and the Periodic Law

2.1 In response to questions DEMONSTRATE an understanding of the structure of the atom and its relationship to the periodic law to a 60% level of competence.

2.1.1 WRITE, from memory, chemical symbols for the first 18 elements.

2.1.2 DESCRIBE the composition of the atom in terms of protons, neutrons, and electrons.

2.1.3 DESCRIBE the approximate size, relative mass and charge on an atom, proton, neutron and electron.

2.1.4 WRITE the symbols for and the charges on positive and negative ions.

2.1.5 DEFINE atomic number.

2.1.6 DEFINE atomic weight.

2.1.7 RELATE atomic structure to groups and families.

2.1.8 EXPLAIN the Bohr model of the atom in terms of energy levels and electronic structure.

2.1.9 RELATE s, p, d sub-orbitals to energy levels.

2.1.10 DESCRIBE shapes of s, p, d, sub-orbitals, including their relative distances from the nucleus.

2.1.11 RELATE principle quantum numbers (n) to the periodic table.

2.1.12 RELATE azimuthal quantum numbers (ℓ) to principle quantum numbers.

2.1.13 RELATE magnetic or orientational quantum numbers to azimuthal quantum numbers.

2.1.14 DESCRIBE the filling order of atomic orbitals and sub-orbitals.

2.1.15 RELATE spin quantum numbers to sub-orbitals.

UNIT 3.0 <u>The Relationship of Periodic Classification to the Properties of Elements</u>.

3.1 In response to questions, given specific examples and a periodic table, DEMONSTRATE knowledge of the relationships between periodic classification of elements and the properties of the elements at a 60% competency level.

3.1.1 APPLY the concept of Avogadro's number to elements of different masses.

3.1.2 DEFINE groups of the periodic table in terms of electron configuration.

3.1.3 DEFINE periods of the periodic table in terms of electron configuration.

3.1.4 IDENTIFY locations of activite metals, transition elements and nonmetals on the periodic table.

3.1.5 DESCRIBE variations in ionization energies based upon elements positions on the periodic table.

3.1.6 DESCRIBE the relationship between periodic classification of elements and their atomic radii.

3.1.7 EXPLAIN the concept of electron affinity and its relationship to ionization energies.

3.1.8 DESCRIBE the relationship between melting point and periodic classification.

3.2 Given specific questions, IDENTIFY at least two uses of the periodic table.

3.2.1 DESCRIBE the relationship between compound formation and atomic number.

3.2.2 PREDICT diatomic molecule formation based upon atomic number.

UNIT 4.0 General Principles of Chemical Bonding

4.1 Given a periodic table, APPLY oxidation numbers to predict chemical compound formation between specific elements to a 60% level of competence.

4.1.1 DETERMINE the number of valence electrons of atoms in a compound and, from this, write a Lewis structure.

4.1.2 WRITE electron configurations for ions.

4.1.3 PREDICT the common oxidation states of metalic and nonmetalic elements from their positions on the periodic table.

4.1.4 ASSIGN oxidation numbers to ions and molecules.

4.1.5 ASSIGN oxidation numbers to atoms in molecules and ions.

4.1.6 DESCRIBE the difference between valance numbers and oxidation states.

4.2 Given table of electronegativities, DESCRIBE the concept of electronegativity to a 60% level of competence.

4.2.1 EXPLAIN the significance of electronegativity.

4.2.2 RELATE the electronegativity of an element to its position on the periodic table.

4.2.3 PREDICT bond polarities of diatomic compounds using a table of electronegativities.

4.2.4 COMPARE compounds based upon electronegativities to determine which bonds have more covalent and which bonds have more ionic character.

4.3 Given a periodic table, DESCRIBE ionic, metalic, and covalent bonding to the 60% competency level.

4.3.1 PREDICT ionic and covalent bonds using the periodic table.

4.3.2 DESCRIBE the ionic-covalent bond continuum in terms of time-sharing of bonding orbital electrons.

4.3.3 WRITE Lewis structures for compounds containing covalent bonds.

4.3.4 DESCRIBE bonding sp, sp^2 and sp^3 orbital shapes.

4.3.5 RELATE free electrons to electron clouds and metalic bonds.

UNIT 5.0 Chemical Symbols, Formulae and Equations.

5.1 Given specific chemical equations, DESCRIBE four general types of chemical changes.

5.1.1 IDENTIFY replacement reactions.

5.1.2 IDENTIFY rearrangement reactions.

5.1.3 IDENTIFY double replacement reactions.

5.1.4 IDENTIFY combination reactions.

5.1.5 DERIVE formulae for replacement, rearrangement and combination reactions.

5.2 Given specific problems, UNDERSTAND the significance of Avogadro's number, moles, formula weights and molecular weights in balancing chemical equations 60% of the time.

5.2.1 BALANCE replacement, rearrangement and combination reactions using Avogadro's number.

5.2.2 BALANCE replacement, rearrangement and combination reactions using moles.

5.2.3 DEFINE formula weights.

5.2.4 BALANCE replacement, rearrangement and combination reactions using molecular weights.

5.3 Given specific problems, UNDERSTAND percent composition based upon chemical formulae 60% of the time.

5.3.1 SOLVE problems based upon percent composition .

5.4 Given specific real-life situations, APPLY balanced chemical equations to obtain solutions to the 60% competency level.

5.4.1 MAKE calculations based upon balanced chemical equations.

UNIT 6.0 The Gaseous State and Kinetic Molecular Theory.

6.1 Given specific questions, DEMONSTRATE an understanding of the gaseous state and kinetic molecular theory with 60% competency.

6.1.1 LIST the assumptions upon which the kinetic theory of gases is based.

6.1.2 EXPLAIN the assumptions on which the kinetic theory of gases is based.

6.1.3 COMPARE the general characteristics of gases to other states of matter.

6.1.4 LIST ways in which gases are distinctly different from other states of matter.

6.1.5 LIST the variables required to define a substance as a gas.

6.1.6 DEFINE standard temperature and pressure (STP).

6.1.7 CORRECT for the effects of water vapor pressure in calculating the quantity of a gas collected over water.

6.1.8 EXPLAIN the concept of gas density.

6.1.9 DESCRIBE how gas density is related to temperature pressure and molecular weight.

6.1.10 APPLY Boyle's law to the solution of problems.

6.1.11 APPLY Charle's law to the solution of problems.

6.1.12 DESCRIBE Graham's law.

6.1.13 APPLY Avogadro's hypothesis to the solution of problems.

6.1.14 CALCULATE molecular weight, given gas density, under defined conditions of temperature and pressure.

6.1.15 CALCULATE gas density under stated conditions, given molecular weight.

6.1.16 DESCRIBE how gas molecules are distributed over a range of speeds.

6.1.17 DESCRIBE how gas molecule distribution changes with temperature.

6.1.18 EXPLAIN the correction terms to P and V that appear in the Van der Walls equation of state for a gas.

6.1.19 EXPLAIN the concept of mean free path.

6.1.20 EXPLAIN how the concept of mean free path relates to the diffusion rates of molecules in the gaseous state.

6.1.21 DESCRIBE the relationship between mean free path and thermal conductivity.

6.2 Given specific situations, DEMONSTRATE an understanding of gram-molecular volume of gases in terms of molecular, atomic and equivalent weights to the 60% competency level.

6.2.1 EXPLAIN the way in which pressure, volume and temperature are related in the ideal gas equation.

6.2.2 EXPLAIN the origin of deviations shown by real gases from the relationship PV/RT=1 for a mole of ideal gas.

6.2.3 LIST two factors responsible for deviations of a gas from ideal behavior.

6.2.4 DERIVE atomic weights from specific heats of gases.

6.2.5 DESCRIBE the relationship between atomic weights and equivalent weights.

UNIT 7.0 The Liquid and Solid States

7.1 In general and specifically, DEMONSTRATE knowledge of the differences between liquids and solids and how they react under different conditions to the 60% competency level.

7.1.1 DISTINGUISH between liquids and solids on a molecular level.

7.1.2 EXPLAIN the properties of each phase (surface tension, viscosity, vapor pressure, boiling point, melting point) using kinetic-molecular theory and the concept of inter-molecular attractions.

7.1.3 DISTINGUISH between crystalline and amorphos solids.

7.1.4 EXPLAIN the nature of equilibria that may exist between phases.

7.1.5 ACCOUNT for enthalpy changes that accompany phase changes.

7.1.6 DESCRIBE the relationship between pressure, vapor pressure, temperature and boiling point.

7.1.7 DEFINE critical temperature in terms of liquid formation.

7.1.8 DEFINE critical pressure in terms of liquid formation.

7.1.9 ACCOUNT for differences in critical temperature and pressure of different substances based upon intermolecular forces.

7.1.10 CALCULATE the density of a substance from the number of particles per unit cell.

7.1.11 EXPLAIN the various types of intermolecular forces.

7.1.12 DRAW phase diagrams.

7.1.13 USE phase diagrams to predict phases present at specific temperatures and pressures.

UNIT 8.0 Solutions

8.1 Given specific situations, DEMONSTRATE an understanding of the definitions of solutions at or above a 60% level of competence.

8.1.1 DEFINE gas-liquid solutions.

8.1.2 DEFINE liquid-liquid solutions.

8.1.3 DEFINE solid-liquid solutions.

8.1.4 DEFINE solvent.

8.1.5 DEFINE solute.

8.1.6 CALCULATE concentration in molarity, molality, normality, mole fraction and weight percent.

8.1.7 CONVERT concentrations in one type unit to another type unit (e.g. molarity → normality, etc.)

8.1.8 DESCRIBE dilute solutions.

8.1.9 DESCRIBE concentrated solutions.

8.1.10 DEFINE unsaturated, saturated, and supersaturated solutions.

8.1.11 DESCRIBE the difference between a colloid and a true solution.

8.2 Given specific situations, DEMONSTRATE knowledge of the effects of solutes upon the properties of the solvent, at or above the 60% level of competence.

8.2.1 DESCRIBE the solution process including bonds made and broken when a substance dissolves.

8.2.2 Explain energy changes occurring in the solution process in terms of solvent-solute forces.

8.2.3 APPLY LaChatelier's principle to account for the effects of temperature and pressure upon solubilities.

8.2.4 DESCRIBE the effect of solute concentration upon solvent vapor pressure.

8.3 Given specific conditions (solvents and solutes) show an understanding of the meaning of solutions of electrolytes at at least the 60% competency level.

8.3.1 PREDICT whether a substance is or is not an electrolyte.

8.3.2 DETERMINE whether substances are strong or weak electrolytes.

8.3.3 EXPLAIN the difference between the magnitude of changes in colligative properties caused by electrolytes as compared to those caused by nonelectrolytes.

8.3.4 IDENTIFY the properties of ionic compounds and solutions.

8.3.5 DESCRIBE the dissolution mechanism for ionic compounds.

UNIT 9.0 Oxidation-Reduction Reactions

9.1 Given a list of oxidation-reduction reaction related terms, KNOW their definitions with 60% competence.

9.1.1 DEFINE oxidation.

9.1.2 DEFINE reduction.

9.1.3 DEFINE oxidizing agent.

9.1.4 DEFINE reducing agent.

9.1.5 IDENTIFY the oxidizing agent.

9.1.6 IDENTIFY the reducing agent.

9.1.7 RECOGNIZE redox reactions.

9.1.8 EXPLAIN how gram-equivalent weights relate to oxidizing and reducing agents.

9.2 Given unbalanced oxidation-reduction (redox) equations, UNDERSTAND how to balance those equations with 60% competence.

9.2.1 BALANCE redox equations by using the change in oxidation number method.

9.2.2 BALANCE redox equations by the ion-electron method.

9.2.3 BALANCE redox reactions using the half reaction method.

9.2.4 DESCRIBE the steps used in balancing redox reactions by the algebraic method.

UNIT 10.0 Acids, Bases and Salts

10.1 Given specific and general acid-base situations, KNOW the properties of acids, bases and salts, at 60% competency level.

10.1.1 DESCRIBE the effect of an acid on a solvent.

10.1.2 DESCRIBE the effect of a base on a solvent.

10.1.3 IDENTIFY the source of a salt in an acid base reaction.

10.1.4 EXPLAIN the meaning of acid-base neutralization.

10.1.5 EXPLAIN the process that occurs when an acid dissolves in water.

10.1.6 DESCRIBE the forms in which a proton exists in water.

10.1.7 DEFINE autoionization.

10.1.8 WRITE an ion product constant for autoionization.

10.1.9 DEFINE pH.

10.1.10 CALCULATE pH for a given OH^- .

10.1.11 DEFINE pOH.

10.1.12 CALCULATE pOH for a given H^{+-} .

10.1.13 DEFINE strong acid.

10.1.14 DEFINE strong base.

10.1.15 CALCULATE Ka.

10.1.16 CALCULATE Kb.

10.1.17 WRITE the acid-dissociation constant expression for a weak acid in water.

10.1.18 WRITE the base dissociation constant for a weak base in water.

10.1.19 DEFINE conjugate acid.

10.1.20 DEFINE conjugate base.

10.2 Given a list of terms, KNOW how the terms apply to acid-base theories at or above the 60% competency level.

10.2.1 DEFINE acid, base, conjugate acid and conjugate base in terms of the Bronstead-Lowry theory.

10.2.2 DEFINE acid and base in terms of the Lewis theory.

UNIT 11.0 Chemical Equilibrium

11.1 Given general reaction conditions, DEMONSTRATE an understanding of the factors affecting reaction rates at or above the 60% competency level.

11.1.1 EXPLAIN how concentration affects reaction rate.

11.1.2 DESCRIBE how temperature affects reaction rate.

11.1.3 EXPLAIN the collision theory of reaction rates.

11.1.4 DESCRIBE the meaning of slowest step in reaction mechanisms.

11.1.5 DEFINE calalyst.

11.1.6 EXPLAIN how pressure affects reaction rates.

11.2 Given general and specific reaction conditions, DEMONSTRATE an understanding of the meaning of equilibrium constants with 60% competency.

11.2.1 EXPLAIN how LeChatelier's principle is related to chemical equilibrium.

11.2.2 WRITE the equilibrium constant expression for a balanced chemical equation.

11.2.3 CALCULATE Kc.

11.2.4 CALCULATE Kp.

11.2.5 CALCULATE the reaction quotient, Q.

11.2.6 CALCULATE equilibrium concentrations.

UNIT 12.4 Ionic Equlibria and Weak Electrolytes

12.1 Using specific situations, DEMONSTRATE an understanding of the meaning of ionic equilibria at or above the 60% competency level.

12.1.1 DEFINE common ion effect.

12.1.2 CALCULATE the effect of an added common ion on the solubility of a slightly soluble salts.

12.1.3 DESCRIBE how complex formation can affect the solubility of slightly soluble salts.

12.1.4 EXPLAIN the effect of pH on a solubility equilibrium involving a basic or acidic ion.

12.1.5 CALCULATE ion product constants of weak acids and bases.

12.1.6 CALCULATE pH at any point on a titration curve.

12.1.7 DESCRIBE how a buffer solution operates to control pH.

12.1.8 DESCRIBE titration curves for titration of weak acid by a strong base and a weak base by a strong acid.

12.1.9 DESCRIBE how a buffer solution is made.

UNIT 13.0 The Solubility Product Principle

13.1 Given specific data, DEMONSTRATE an understanding of the solubility product principle at or above the 60%

13.1.1 CALCULATE K_{sp} from molar solubilities.

13.1.2 CALCULATE molar solubilities from K_{sp}.

13.1.3 MAKE calculations involving precipitation of slightly soluble electrolytes.

UNIT 14.0 Electrochemistry

14.1 Given terms, diagrams and problems, DEMONSTRATE a knowledge of the basic principles of electrochemistry with 60% competence.

14.1.1 DEFINE fundamental electrical units.

14.1.2 DIAGRAM simple voltaic and electrolytic cells.

14.1.3 DESCRIBE Faraday's law and how it relates to electrolysis.

14.1.4 CALCULATE the EMF generated by a voltaic cell.

14.1.5 CALCULATE the minimum EMF required to cause an electrolytic cell reaction to proceed.

14.1.6 USE electrode potentials to predict whether a reaction will be spontaneous.

14.1.7 CALCULATE EMF's using the Nernst equation.

14.1.8 CALCULATE ion concentrations using Nernst equation.

UNIT 15.0 Chemistry and Properties of Liquified Natural Gas

15.1 In response to general questions, DEMONSTRATE an understanding of the properties of liquid natural gas with 60% competence.

15.1.1 DESCRIBE the composition of natural gas.

15.1.2 DESCRIBE the liquid ratio of liquid natural gas.

UNIT 16.0 Alloys

16.1 Given a list of terms, DEMONSTRATE an understanding of the definitions of alloys.

16.1.1 DESCRIBE what is meant by alloy.

16.1.2 DEFINE solution alloy.

16.1.3 DEFINE substitutional alloy.

16.1.4 DEFINE interstitial alloy.

16.1.5 DEFINE heterogeneous alloy.

16.1.6 DEFINE intermetallic alloy.

UNIT 17.0 Nuclear Chemistry

17.1 Given specific conditions, DEMONSTRATE an understanding of the meaning of nuclear chemistry with 60% competence.

17.1.1 WRITE nuclear symbols for protons, neutrons, alpha particals and positrons.

17.1.2 DETERMINE the effect of different types of decay on the proton-neutron ratio.

17.1.3 PREDICT the type of decay that a nucleus will undergo based on its position relatives to the belt of stability.

17.1.4 BALANCE nuclear equations.

17.1.5 USE the half-life of a substance to predict the amount of radioisotope present after a given period.

17.1.6 EXPLAIN the use of radioisotopes in dating.

17.1.7 EXPLAIN detection of radioactivity.

17.1.8 DESCRIBE the basic design of a Geiger counter.

17.1.9 CALCULATE energy or mass change in a reaction using $\Delta E = mc^2$

17.1.10 CALCULATE binding energies in the nucleus.

17.1.11 DEFINE units used to describe radioactivity.

17.1.12 EXPLAIN fission.

17.1.13 EXPLAIN fusion.

UNIT 18.0 Corrosion

18.1 Given specific questions and diagrams, DEMONSTRATE an understanding of the corrosion process, including corrosion cells to a 60% competency level.

18.1.1 DEFINE anode and cathode.

18.1.2 DESCRIBE the reactions that take place at the anode.

18.1.3 DESCRIBE the reactions that take place at the cathode.

18.1.4 DIAGRAM corrosion cells without cathode protection.

18.1.5 EXPLAIN the principles underlying cathode protection.

18.1.6 DIAGRAM corrosion cells using cathode protection.

18.1.7 CALCULATE metal loss caused by corrosion.

18.1.8 IDENTIFY causes of corrosion.

18.1.9 DESCRIBE the results of corrosion.

18.1.10 LIST chemical methods of protection against corrosion.

18.1.11 DEFINE electrolysis.

18.1.12 LIST the types of corrosion which occur in marine power plants.

18.1.13 DESCRIBE the causes of specific corrosion reactions that occur in steam power plants.

18.1.14 DESCRIBE how protective coatings and anode-cathode sequences may be used to protect ship hulls.

18.1.15 IDENTIFY the relationship between half reaction potentials and corrosion.